TOPIC

3

Stoichiometry: CALCULATIONS WITH CHEMICAL FORMULAS AND EQUATIONS

The content in this topic is the basis for mastering Learning Objectives 1.2, 1.3, 1.4, 1.17, 1.18, 3.1, and 3.4 as found in the Curriculum Framework.

You may have learned some of this material in first-year chemistry. When you finish reviewing this topic, be sure you are able to:

- Balance a chemical equation using symbols for atoms and molecules and particle drawings
- Use the mole as a quantitative model for chemical composition
- Convert moles to mass, number of particles, and volume of a gas
- Apply mathematical models to explain the composition of compounds
- Calculate the percentage composition of a compound
- Calculate the empirical and molecular formulas of a compound and of a hydrate from combustion and decomposition data
- Apply mathematical calculations to mass data to infer the identity of a substance and/or its purity
- Calculate the masses and moles of reactants and products using stoichiometry
- Determine limiting reactants and percent yields from experimental data

Section 3.1 Chemical Equations

Stoichiometry is the area of study that examines the quantities of substances involved in chemical reactions.

A **chemical reaction** is a process by which one or more substances are converted to other substances.

Chemical equations use chemical formulas to symbolically represent chemical reactions. For example, the chemical sentence, "Hydrogen burns in air to produce water," can be expressed as chemical symbols or as a molecular picture of the particles as shown in Figure 3.1.

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Reactants		Products
$2\,H_2 + O_2$	$\longrightarrow$	$2\,H_2O$

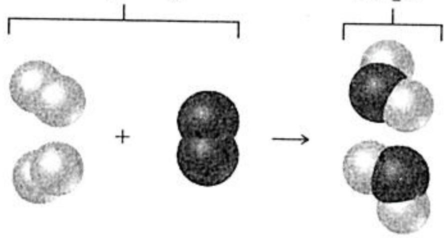

Figure 3.1 A balanced chemical equation.

When the reaction is more complex, chemists prefer to write the equation using only chemical symbols to represent the particle pictures. For example, the following chemical equation describes how butane, C_4H_{10}, burns in air:

$$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(l)$$

$$\text{reactants} \rightarrow \text{products}$$

The equation is a "chemical sentence" written in the symbolic "words" of chemical formulas and symbols. The formulas on the left are reactants, and the formulas on the right are products. The equation reads:

"Two molecules of gaseous butane react with thirteen molecules of oxygen gas to produce eight molecules of carbon dioxide gas and ten molecules of liquid water."

A **balanced chemical equation** has an equal number of atoms of each element on each side of the arrow. The coefficients preceding each formula "balance" the equation. Notice that, because of the coefficients, on each side of the arrow there are eight carbon atoms, twenty hydrogen atoms, and twenty-six oxygen atoms.

The symbols (g), (l), (s), and (aq) are used to designate the physical state of each reactant and product: gas, liquid, solid, and aqueous (dissolved in water).

Your Turn 3.1

Write a chemical sentence to illustrate how to "read" the following chemical equation:

$$CO_2(g) + 2NH_3(g) + H_2O(l) \rightarrow 2NH_4^{+}(aq) + CO_3^{2-}(aq)$$

Convert each formula of the equation to a particle representation. Write your answer in the space provided.

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To balance simple equations, start by balancing atoms other than hydrogen or oxygen. Balance hydrogen atoms next to last and balance oxygen atoms last.

Example:

Balance the following equation:

$H_2SO_4(aq) + NaHCO_3(s) \rightarrow CO_2(g) + H_2O(l) + Na_2SO_4(aq)$

Solution:

1. S and C are already balanced so start with Na.

$H_2SO_4(aq) + 2NaHCO_3(s) \rightarrow CO_2(g) + H_2O(l) + Na_2SO_4(aq)$

2. Now balance carbon.

$H_2SO_4(aq) + 2NaHCO_3(s) \rightarrow 2CO_2(g) + H_2O(l) + Na_2SO_4(aq)$

3. Next balance H.

$H_2SO_4(aq) + 2NaHCO_3(s) \rightarrow 2CO_2(g) + 2H_2O(l) + Na_2SO_4(aq)$

4. Check to see that oxygen is balanced and double-check all other atoms. (As in algebraic equations, the numeral 1 is usually not written.)

Section 3.2 Some Simple Patterns of Reactivity

Predicting the products of chemical reactions is an essential skill to acquire in the study of chemistry. Sometimes reactions fall into simple patterns and recognizing these patterns can be helpful in predicting which products will be produced from given reactants. Topic 4 addresses predicting products of chemical reactions in more depth. For now, here are a few patterns to learn to recognize.

1. In a combination reaction, two elements combine to form one compound. (Other combinations are possible and Topic 4 describes better ways to predict what will happen.)

Example:

A metal reacts with a nonmetal to produce an ionic compound. Write an equation to describe what happens when solid sodium is exposed to chlorine gas.

Solution:

$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$

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Your Turn 3.2

Write an equation to describe what happens when solid magnesium metal reacts at high temperature with nitrogen gas. Write your answer in the space provided.

2. In a decomposition reaction, one reactant changes to two or more products.

Example:

Upon heating, metal carbonates decompose to yield metal oxides and carbon dioxide. Write an equation to describe what happens when solid magnesium carbonate is heated.

Solution:

$MgCO_3(s) \rightarrow MgO(s) + CO_2(g)$

Your Turn 3.3

Write an equation to describe what happens when liquid water is decomposed to its elements by an electrical current. Write your answer in the space provided.

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3. A combustion reaction usually involves oxygen, often from air, reacting with hydrocarbons or other organic molecules containing carbon, hydrogen, and oxygen to produce carbon dioxide and water.

Example:

Write an equation to describe what happens when liquid ethanol burns in air.

Solution:

$CH_3CH_2OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$

(Recall that encoded in the name "ethanol" is a two-carbon alcohol having the –OH functional group.)

Your Turn 3.4

Write an equation to describe what happens when liquid hexane is burned in air. Write your answer in the space provided.

Section 3.4 Avogadro's Number and the Mole

Avogadro's number is 6.02×10^{23}. It represents the number of atoms in exactly 12 g of isotopically pure ^{12}C. Because all atomic masses are based on ^{12}C, the atomic mass of any element expressed in grams represents Avogadro's number (6.02×10^{23}) of atoms of that element.

A **mole** is the amount of matter that contains 6.02×10^{23} atoms, ions, molecules, or formula units.

The **molar mass** of a substance is the mass in grams of one mole of that substance. To calculate a molar mass of any substance, add the atomic masses of all the atoms in its formula. (For convenience, atomic masses are often rounded to three significance figures.) Table 3.1 shows the molar masses of various substances.

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Common misconception: "Molar mass" is a universal term that is often used to replace the terms atomic mass, molecular mass, and formula mass. Molar mass is used to express the mass of one mole of any substance, whether it is an atom (atomic mass), a molecule (molecular mass), an ion (formula mass), or an ionic compound (formula mass).

Table 3.1 The molar mass of any substance is the mass in grams of 6.02×10^{23} particles or one mole of that substance.

Formula	Number of Particles	Representative Particles	Molar Mass	Alternate Term
Ar	6.02×10^{23}	Atoms	39.9 g/mol	Atomic mass
CO_2	6.02×10^{23}	Molecules	44.0 g/mol	Molecular mass
NaBr	6.02×10^{23}	Formula units	103 g/mol	Formula mass
CO_3^{2-}	6.02×10^{23}	Ions	60.0 g/mol	Formula mass

Grams, moles, and representative particles (atoms, molecules, ions, or formula units) are converted from one to another using the "Mole Road" described in Figure 3.2.

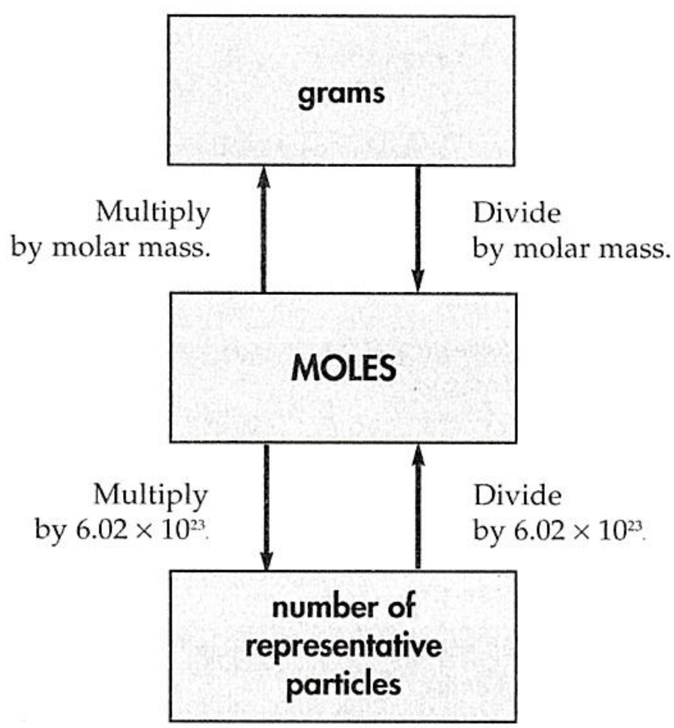

Figure 3.2 The "Mole Road." Divide to convert to moles. Multiply to convert from moles.

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Calculating Percentage Composition of a Compound

The percentage composition of a compound is the percentage by mass contributed by each element in the compound. To calculate the percentage composition of an element in any formula, divide the molar mass of the element multiplied by the number of times it appears in the formula, by the molar mass of the formula, and multiply by 100.

% composition = 100 × (molar mass of element × subscript for element)/(molar mass of substance)

Example:

What is the percentage composition of Na_2CO_3?

Solution:

$$\% \, Na = 100 \times 2(23.0) \, g \, Na/[2(23.0) + 12.0 + 3(16.0) \, g] = 43.4\% \, Na$$

$$\% \, C = 100 \times 12.0 \, g \, C/[2(23.0) + 12.0 + 3(16.0) \, g] = 11.3\% \, C$$

$$\% \, O = 100 \times 3(16.0) \, g/[2(23.0) + 12.0 + 3(16.0) \, g] = 45.3\% \, O$$

Section 3.5 Empirical Formulas from Analyses

The **empirical formula** for a compound expresses the simplest ratio of atoms in the formula. The percentage composition of a compound can be determined experimentally by chemical analysis and the empirical formula can be calculated from the percentage composition.

Example:

What is the empirical formula of a compound containing 68.4% chromium and 31.6% oxygen?

Solution:

In chemistry, percentage always means mass percentage, unless specified otherwise. The data mean that for every 100 g of compound, there are 68.4 g of Cr and 31.6 g of O.

1. *Write the formula using number of grams.*

$Cr_{68.4} O_{31.6}$

2. *Convert grams to moles by dividing by the molar mass of each element.*

$Cr_{68.4/52.0} O_{31.6/16.0} = Cr_{1.315} O_{1.975}$

3. *Convert to small numbers by dividing each mole quantity by the smaller mole quantity.*

$Cr_{1.315/1.315} O_{1.975/1.315} = Cr_1 O_{1.5}$

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4. *If necessary, multiply each mole quantity by a small whole number that converts all quantities to whole numbers.*

$Cr_{1\times2}O_{1.5\times2} = Cr_2O_3$

Molecular Formulas from Empirical Formulas

A **molecular formula** tells exactly how many atoms are in one molecule of the compound. The subscripts in a molecular formula are always whole number multiples of the subscripts in the empirical formula. Molecular formulas can be determined from empirical formulas if the molar mass of the compound is known.

Example:

A compound containing only carbon, hydrogen, and oxygen is 63.16% C and 8.77% H. Mass spectrometry shows that the compound has a molar mass of 114 g/mol. What is its empirical formula and its molecular formula?

Solution:

1. *Write the formula using number of grams.*

$C_{63.16}\,H_{8.77}\,O_{28.07}$

(Calculate the grams of oxygen by subtracting the grams of carbon and grams of hydrogen from 100 *g.*

$x\ g\ O = 100\ g - 63.16\ g\ C - 8.77\ g\ H = 28.07\ g\ O.)$

2. *Convert grams to moles by dividing by the molar mass of each element.*

$C_{63.16/12.0}\,H_{8.77/1.00}\,O_{28.07/16.0} = C_{5.263}\,H_{8.77}\,O_{1.754}$

3. *Convert to small numbers by dividing each mole quantity by the smallest mole quantity.*

$C_{5.263/1.754}\,H_{8.77/1.754}\,O_{1.754/1.754} = C_3\,H_5\,O_1$

empirical formula

4. *All quantities are whole numbers.*

5. *Divide the known molar mass by the mass of one mole of the empirical formula. The result produces the integer by which you multiply the empirical formula to obtain the molecular formula.*

$(114\ g/mol)/(57.0\ g/mol) = 2$

$C_{3\times2}\,H_{5\times2}\,O_{1\times2} = C_6\,H_{10}\,O_2$

molecular formula

Empirical Formulas from Combustion Analysis

When a compound containing carbon, hydrogen, and oxygen is completely combusted, all the carbon is converted to carbon dioxide, and all the hydrogen becomes water. The empirical formula of the compound can be calculated from the measured masses of the products.

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Example:

A 3.489 g sample of a compound containing C, H, and O yields 7.832 g of CO_2 and 1.922 g of water upon combustion. What is the simplest formula of the compound?
The unbalanced chemical equation is:

$$C_xH_yO_z + O_2 \rightarrow CO_2 + H_2O$$

X is the number of moles of carbon in the compound because all the carbon in the compound is converted to CO_2. X equals the number of moles of CO_2 because there is one mole of carbon in one mole of CO_2.

$$X = mol\ C = 7.832\ g\ CO_2/44.0\ g/mol = 0.178\ mol\ C$$

Y is the number of moles of hydrogen in the compound because all the hydrogen becomes water. Y equals twice the number of moles of water because there are two moles of hydrogen in one mole of water.

$$Y = mol\ H = (1.922\ g\ H_2O/18.0\ g/mol) \times 2 = 0.2136\ mol\ H$$

Z is the number of mol of O. To obtain the number of grams of O in the compound, subtract the number of grams of C ($X\ mol \times 12.0\ g/mol$) and the number of grams of H ($Y\ mol \times 1.00\ g/mol$) from the total grams of the compound. Convert the result to moles of O by dividing by 16.0 g/mol.

$$Z = mol\ O = [3.489 - (12.0)(0.178) - (1.00)(0.2136)]/16.0$$
$$= 0.0712\ mol\ O$$

Convert to small numbers by dividing each mole quantity by the smallest mole quantity.

$$C_{0.178}\ H_{0.2136}\ O_{0.0712} =$$
$$C_{0.178/0.0712}\ H_{0.2136/0.0712}\ O_{0.0712/0.0712} = C_{2.5}H_3O_1$$

Finally, multiply each mole quantity by a small whole number that converts all quantities to whole numbers.

$$C_{2.5\times2}\ H_{3\times2}\ O_{1\times2} = C_5H_6O_2$$

Formulas of Hydrates

Ionic compounds often form crystal structures called hydrates by acquiring one or more water molecules per formula unit. For example, solid sodium thiosulfate decahydrate, $Na_2S_2O_3 \cdot 10H_2O$, has ten water molecules per formula unit of sodium thiosulfate. Heating a sample of hydrate causes it to lose water. The number of water molecules per formula unit can be calculated from the mass difference before and after heating.

Example:

When 21.91 g of a hydrate of copper(II) sulfate is heated to drive off the water, 14.00 g of anhydrous copper(II) sulfate remain. What is the formula of the hydrate?

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Solution:

This is a variation of an empirical formula problem. The solution lies in calculating the ratio of H_2O moles to the moles of $CuSO_4$. The chemical equation is:

$$CuSO_4 \cdot H_2O(s) \rightarrow CuSO_4(s) + X\,H_2O(g)$$
$$21.91\ g \qquad\qquad 14.00\ g$$

Calculate the grams of water by subtracting the grams of copper(II) sulfate from the grams of the hydrate.

$x\ g\ H_2O = 21.91\ g - 14.00g = 7.91\ g\ H_2O$

$mol\ H_2O = 7.91\ g\ H_2O/18.0\ g/mol = 0.439\ mol\ H_2O$

$mol\ CuSO_4 = 14.00\ g/159.5\ g/mol = 0.08777\ mol\ CuSO_4$

$mol\ H_2O/mol\ CuSO_4 = 0.439\ mol/0.08777\ mol = 5.00$

The formula has five moles of water per mole of copper(II) sulfate: $CuSO_4 \cdot 5H_2O$.

Quantitative Information from Balanced Equations

Section 3.6

Stoichiometry refers to the quantities of substances involved in chemical reactions. The coefficients in a balanced chemical equation indicate both the relative number of molecules (or formula units) involved in the reaction, and the relative number of moles. For example, the equation for the combustion of octane, C_8H_{18}, a component of gasoline is:

$$2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$$

The four coefficients that balance the equation are proportional to one another and can be used to relate mole quantities of reactants and/or products.

Example:

How many moles of octane will burn in the presence of 37.0 moles of oxygen gas?

Solution:

The balanced equation tells us that two moles of octane will burn in 25 moles of oxygen gas, so the answer to the question involves the ratio 2 mol C_8H_{18} per 25 mol O_2.

$x\ mol\ C_8H_{18} = 37.0\ mol\ O_2\ (2\ mol\ C_8H_{18}/25\ mol\ O_2) = 37.0 \times 2/25$
$= 2.96\ mol\ C_8H_{18}$.

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Alternatively, we can use the mole road:

To convert mol O_2 to mol C_8H_{18}, multiply by $2/25$.

$x\ mol\ C_8H_{18} = 37.0\ mol(2/25) = 2.96\ mol\ C_8H_{18}$

(Always multiply by the coefficient at the head of the arrow and divide by the coefficient at the tail of the arrow.)

$$mol\ C_8H_{18} \underset{\times 25/2}{\overset{\times 2/25}{\rightleftarrows}} mol\ O_2$$

To convert mol C_8H_{18} to mol O_2, multiply by 25/2.

Figure 3.3 shows an expanded mole road to include the relationship between the moles of any reactant or product in a balanced chemical equation.

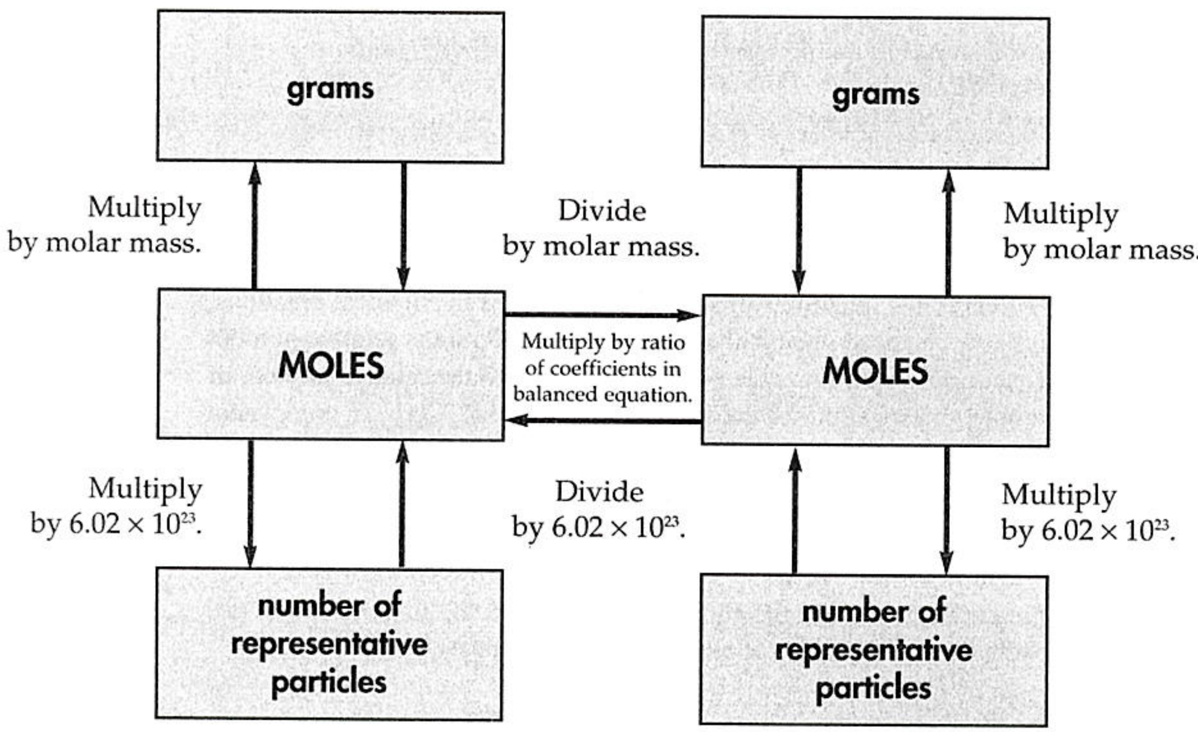

Figure 3.3 The stoichiometry "Mole Road." Divide to convert to moles. Multiply to convert from moles. Multiply by the ratio of coefficients in a balanced equation to convert moles of one substance to moles of another substance.

This stoichiometry mole road allows us to relate any quantity in a balanced equation to any other quantity in the same equation.

Example:

How many grams of carbon dioxide are obtained when 695 g (about 1 gallon) of octane are burned in oxygen?

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Solution:

Follow the road map:

$$2C_8H_{18}(l) + 25O_2(g) \longrightarrow 16CO_2(g) + 18H_2O(g)$$

$695\ g\ C_8H_{18}$ $\qquad\qquad$ **$?\ g\ CO_2$**

Divide by molar mass of octane: **÷ 114 g/mol** $\downarrow$

$\uparrow$ *Multiply by molar mass of CO_2:* **x 44.0 g/mol**

Multiply by ratio of coefficients:

$$mol\ C_8H_{18} \xrightarrow[\textbf{2 mol}]{\textbf{16 mol}} mol\ CO_2$$

x g CO_2 = (695 g)(114 g/mol)(16.0 mol/2 mol)(44.0 g/mol) = 2150 g CO_2

Often chemists use stoichiometric calculations to determine the purity of a substance as a percentage.

Example:

Chalk is composed of a mixture of calcium carbonate and calcium sulfate. To determine the percentage of calcium carbonate in the chalk, a student reacts 100.0 g of chalk with excess hydrochloric acid and finds that the reaction produces 33.0 g of carbon dioxide. What is the percentage of calcium carbonate in the chalk? Assume that calcium sulfate does not react with acid and calcium carbonate reacts according to the following balanced equation:

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

Solution:

The amount of carbon dioxide generated by reaction with the calcium carbonate in the chalk is directly proportional to the amount of calcium carbonate in the mixture.

$x\ g\ CaCO_3 = 33.0\ g\ CO_2(1\ mol\ CO_2/44.0\ g\ CO_2)$
$(1\ mol\ CaCO_3/1 mol\ CO_2)(100.0\ g\ CaCO_3)/1\ mol\ CaCO_3)$
$= 75.0\ g\ CaCO_3$

$\%\ CaCO_3 = (g\ CaCO_3/g\ chalk)(100) = 75.0\ g/100.0\ g = 75\%$

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Section 3.7

Limiting Reactants

The **limiting reactant** is the reactant that is completely consumed in a chemical reaction. The limiting reactant limits the amount of products formed.

The **excess reactant** is usually the other reactant. Some of the excess reactant is left unreacted when the limiting reactant is completely consumed.

A stoichiometric mixture means that both reactants are limiting and both are completely consumed by the reaction. There is an excess of neither. Because of their subtle nature, quantitative limiting reactant problems are among the most difficult. The mole road is a useful tool in solving limiting reactant problems.

Common misconception: Stoichiometry calculations can calculate only how much reactants react or how much products are formed. Stoichiometry cannot calculate how much excess reactant is left unreacted. To calculate how much excess reactant remains after the reaction is complete, first calculate how much is consumed and then subtract that amount from how much total reactant was initially present.

Example:

255 g of octane and 1510 g of oxygen gas are present at the beginning of a reaction that goes to completion and forms carbon dioxide and water according to the following equation.

$$2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O\,(g)$$

a. What is the limiting reactant?

b. How many grams of water are formed when the limiting reactant is completely consumed?

c. How many grams of excess reactant are consumed?

d. How many grams of excess reactant are left unreacted?

Solution:

a. To find the limiting reactant, first compare the number of moles of reactants relative to the ratio in which they react. To do this, divide the number of moles of each reactant by its corresponding coefficient that balances the equation. The resulting lower number always identifies the limiting reactant. Then use the mole road to solve the problem based on the identified limiting reactant.

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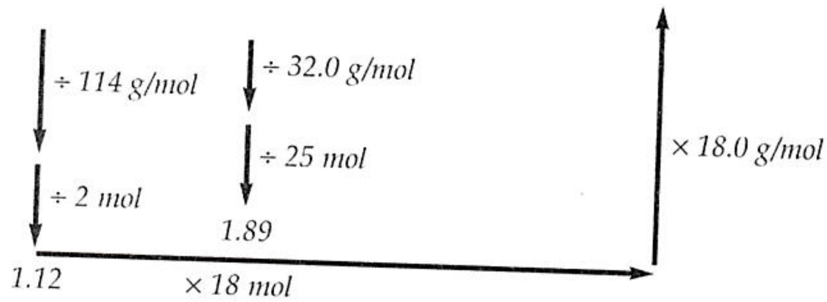
$$2C_8H_{18}(l) \quad + \quad 25O_2(g) \longrightarrow 16CO_2(g) \quad + \quad 18H_2O(g)$$

255 g *1510 g* *? g*

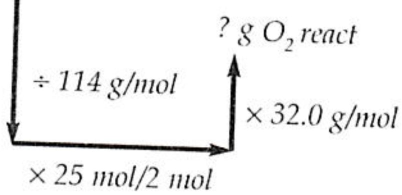


a. C_8H_{18} *is the limiting reactant because* $1.12 < 1.89$.

b. $x\ g\ H_2O = (255\ g/114\ g/mol)(18\ mol/2\ mol)(18.0\ g/mol)$
$= 362\ g\ H_2O$

c. $2C_8H_{18}(l) + 25O_2(g) \longrightarrow 16CO_2(g) + 18H_2O(g)$

255 g *1510 g* *? g*



$x\ g\ O_2 = (255\ g/114\ g/mol)(25\ mol/2\ mol)(32.0\ g/mol)$
$= 895\ g\ O_2\ react$

d. $g\ O_2$ *left unreacted* $= 1510\ g - 895\ g = 615\ g\ O_2$ *unreacted*

Theoretical, Actual, and Percent Yields

The **theoretical yield** of a reaction is the quantity of product that is calculated to form.

The **actual yield** is the amount of product actually obtained and is usually less than the theoretical yield.

The **percent theoretical yield** relates the actual yield to the theoretical yield:

Percent theoretical yield = (actual yield)/(theoretical yield) × 100

Example:

In the previous example, the theoretical yield of water is calculated to be 362 g. What is the percent yield if the actual yield of water is only 312 g?

Solution:

Percent theoretical yield $= 312g/362g \times 100 = 86.2\%$

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Multiple Choice Questions

1. *Ammonia forms when hydrogen gas reacts with nitrogen gas. If equal number of moles of nitrogen and hydrogen are combined, the maximum number of moles of ammonia that could be formed will be equal to*

   A) the number of moles of hydrogen

   B) the number of moles of nitrogen

   C) twice the number of moles of hydrogen

   D) twice the number of moles of nitrogen

   E) two-thirds the number of moles of hydrogen

2. *If $C_4H_{10}O$ undergoes complete combustion, what is the sum of the coefficients when the equation is completed and balanced using smallest whole numbers?*

   A) 8

   B) 16

   C) 22

   D) 25

   E) 32

3. *What are the products when lithium carbonate is heated?*

   A) $LiOH + CO_2$

   B) $Li_2O + CO_2$

   C) $LiO + CO_2$

   D) $LiC + O_2$

   E) $LiO + CO$

4. *Beginning with 48 moles of H_2, how many moles of $Cu(NH_3)_4Cl_2(aq)$ can be obtained if the synthesis of $Cu(NH_3)_4Cl_2(aq)$ is carried out through the following sequential reactions? Assume that a 50% yield of product(s) is (are) obtained in each reaction.*

   1. $3H_2(g) + N_2(g) \rightarrow 2NH_3(g)$

   2. $4NH_3(g) + CuSO_4(aq) \rightarrow Cu(NH_3)_4SO_4(aq)$

   3. $Cu(NH_3)_4SO_4 + 2NaCl \rightarrow Cu(NH_3)_4Cl_2(aq) + Na_2SO_4(aq)$

   A) 1

   B) 2

   C) 4

   D) 8

   E) 12

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5. *What mass of water can be obtained from 4.0 g of H_2 and 16 g of O_2?*

   $2H_2(g) + O_2(g) \rightarrow 2H_2O$

   A) 9.0 g

   B) 18 g

   C) 36 g

   D) 54 g

   E) 72 g

6. *The empirical formula of pyrogallol is C_2H_2O and its molar mass is 126. Its molecular formula is*

   A) C_2H_2O

   B) $C_4H_4O_2$

   C) $C_2H_6O_3$

   D) $C_6H_6O_3$

   E) $C_2H_6O_6$

7. *What is the maximum amount of water that can be prepared from the reaction of 20.0 g of HBr with 20.0 g of $Ca(OH)_2$?*

   $2HBr + Ca(OH)_2 \rightarrow CaBr_2 + 2H_2O$

   A) (20/81) (2/2) (18) g

   B) (20/74) (2/2) (18) g

   C) (20/81) (2/1) (18) g

   D) (20/74) (2/1) (18) g

   E) (20/74) (1/2) (18) g

8. *How many moles of ozone, O_3, could be formed from 96.0 g of oxygen gas, O_2?*

   A) 0.500

   B) 1.00

   C) 2.00

   D) 3.00

   E) 1/16

9. *The percentage of oxygen in $C_8H_{12}O_2$ is*

   A) (16/140) (100)

   B) (32/140) (100)

   C) (16/124) (100)

   D) (140/32) (100)

   E) (32/124) (100)

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10. *A compound contains 48.0% O, 40.0% Ca, and the remainder is C. What is its empirical formula?*

   A) $O_3C_2Ca_2$

   B) O_3CCa_2

   C) O_3CCa

   D) O_3CCa_2

   E) O_2CCa

Free Response Questions

1. *Combustion of 8.652 g of a compound containing C, H, O, and N yields 11.088 g of CO_2, 3.780 g of H_2O, and 3.864 g of NO_2.*

   a. *How many moles of C, H, and N are contained in the sample?*

   b. *How many grams of oxygen are contained in the sample?*

   c. *What is the simplest formula of the compound?*

   d. *If the molar mass of the compound lies between 200 and 300, what is its molecular formula?*

   e. *Write and balance a chemical equation for the combustion of the compound.*

2. *A student finds that 344.0 g of a pure sample of the mineral gypsum contains 80.00 g of calcium.*

   a. *Show a mathematical calculation to demonstrate that the pure mineral sample does not contain pure calcium sulfate.*

   b. *If gypsum is hydrated calcium sulfate, use the data from the experiment to derive the chemical formula for gypsum. (How many waters of hydration does the formula contain?)*